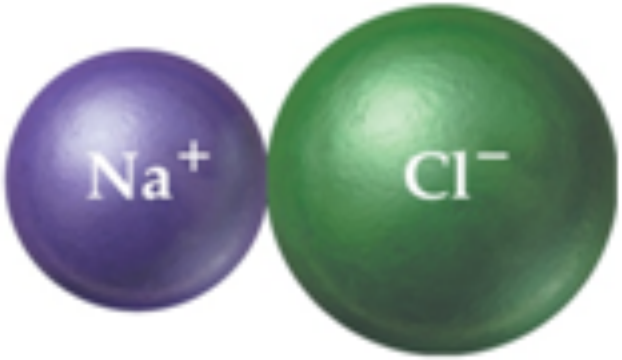
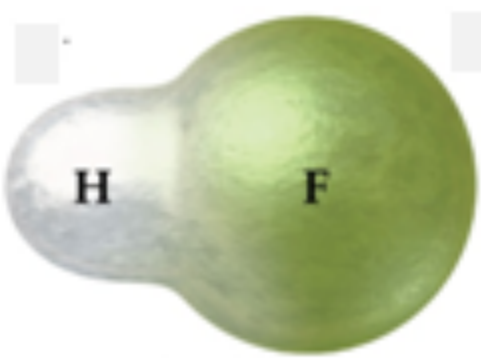
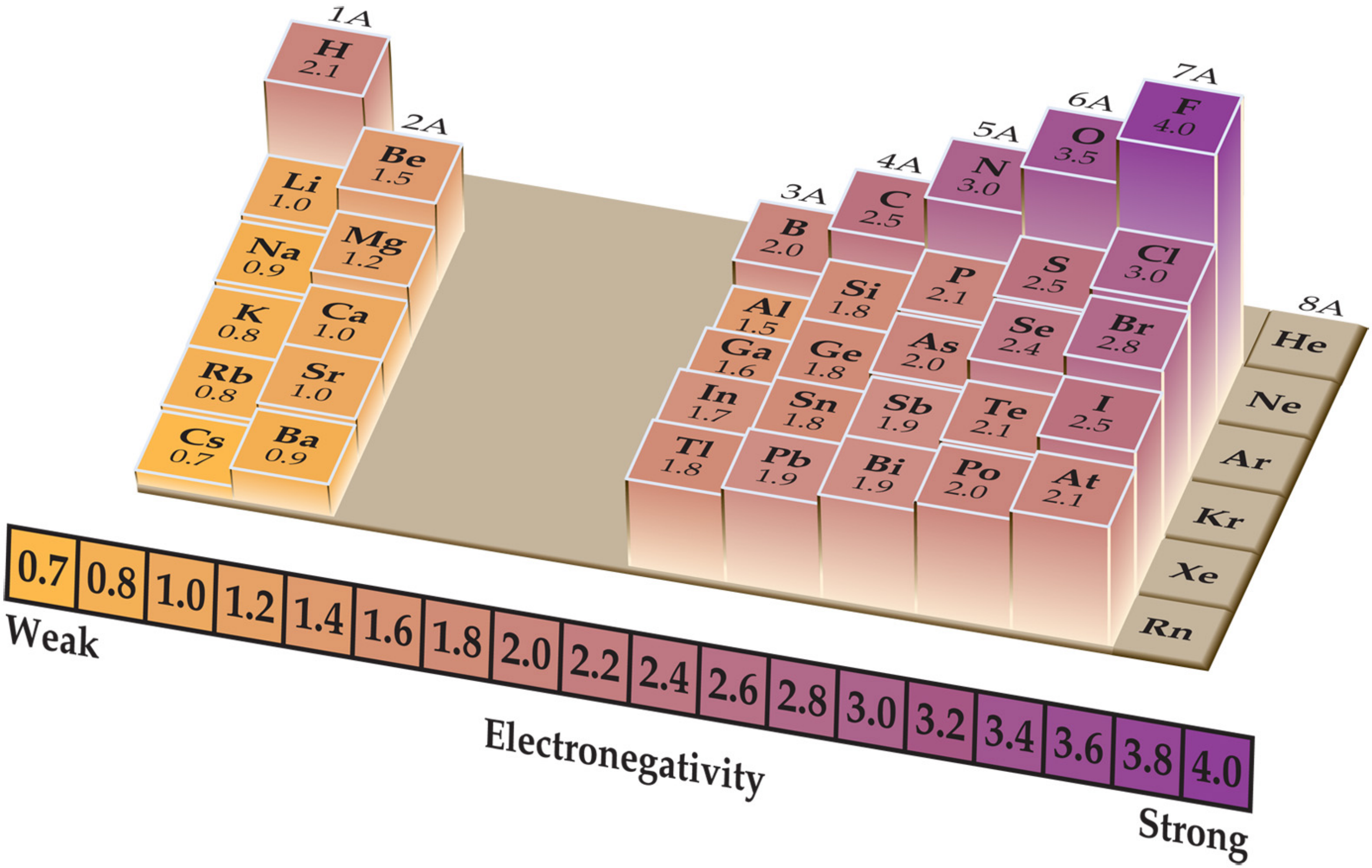


Ionic Bonding

Nonpolar Covalent Bonding

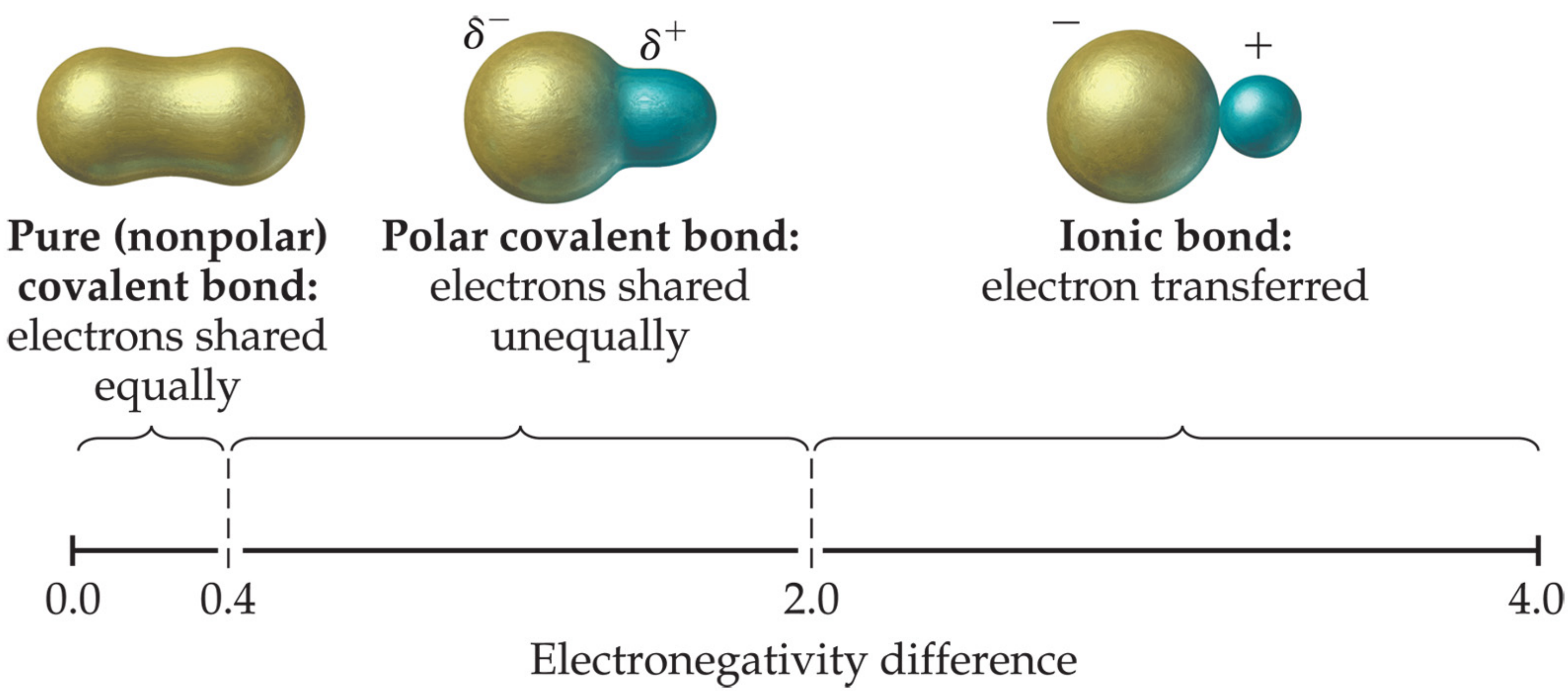


Polar Covalent Bonding


1A
H 2.1
2A
Li 1.0
Be 1.5
Na 0.9
Mg 1.2
K 0.8
Ca 1.0
Rb 0.8
Sr 1.0
Cs 0.7
Ba 0.9
3A
4A
5A
6A
7A
B 2.0
C 2.5
N 3.0
O 3.5
F 4.0
Al 1.5
Si 1.8
P 2.1
S 2.5
Cl 3.0
Ga 1.6
Ge 1.8
As 2.0
Se 2.4
Br 2.8
In 1.7
Sn 1.8
Sb 1.9
Te 2.1
I 2.5
Tl 1.8
Pb 1.9
Bi 1.9
Po 2.0
At 2.1
8A
He
Ne
Ar
Kr
Xe
Rn
0.7 0.8 1.0 1.2 1.4 1.6 1.8 2.0 2.2 2.4 2.6 2.8 3.0 3.2 3.4 3.6 3.8 4.0
Weak
Electronegativity
Strong

Electronegativity and Bond Type

Electronegativity Difference (≤ EN)	Bond Type	Example
zero (0–0.4)	pure covalent	Cl_2
intermediate (0.4–2.0)	polar covalent	HF
large (2.0+)	ionic	NaCl



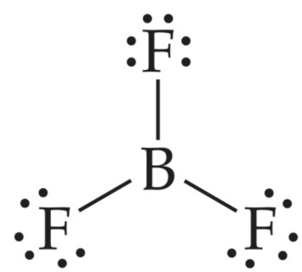
Lewis Structures

1. Write skeletal structure for the molecule.
    - More electropositive atoms in the center, H atoms are always terminal, take into account symmetry.
2. Calculate the total number of electrons by adding the valance electrons of the atoms.
    - For polyatomic ion, add one electron for each negative charge, or subtract one electron for each positive charge.
3. Distribute the electrons among the atoms giving octets (or duets for H) to as many atoms as possible.
    - Begin by placing the bonding electrons, then, give lone pairs to terminal atoms.
4. If any atom lacks an octet, form double or triple bonds as necessary to give them octets.

Exceptions to the Octet Rule

- Molecules and polyatomic ions:
  - containing an odd number of electrons.

  - in which an atom has fewer than an octet of valence electrons.



  - in which an atom has more than an octet of valence electrons.